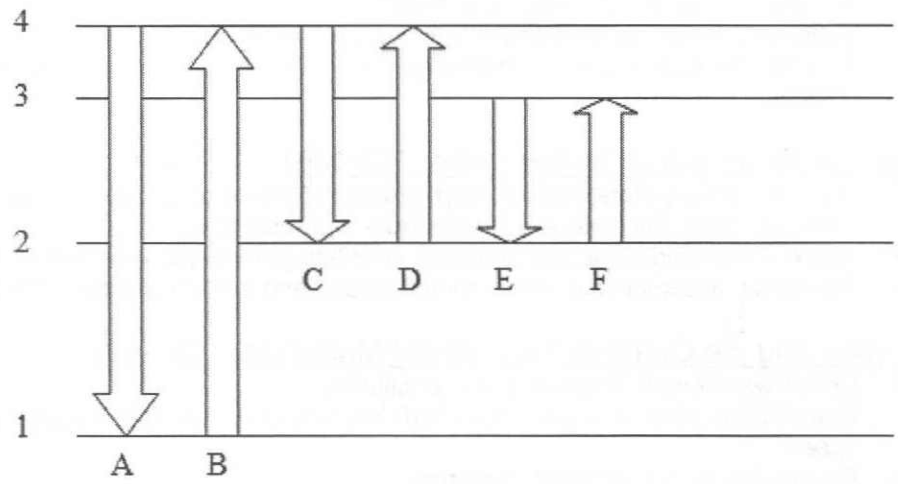
Supplemental Questions

This work will now be due two days before your next test to allow enough time evaluation and return

1. The quantum level occupied by an electron in an atom depends on the energy of the electron. Changes in quantum level are related to absorption or emission of energy. The figure below represents the four lowest energy levels of an atom. (n = 1 to 4). The six lettered arrows represent changes in the energy level of an electron.



   a. Why do these energy levels mean that the atom will show an emission spectrum of discrete lines rather than a continuous spectrum of emitted light?

      Only specific wavelengths are emitted as photons are released in C and E as opposed to all the wavelengths of visible light that are represented in a continuous spectrum

   b. Which three of the lettered energy changes involve absorption of energy by the atom?

      B D F

   c. Which three of the lettered energy changes involve emission of light energy by the atom?

      A C E

   d. Of the three energy changes that involve emission , one results in the emission of blue light, one results in yellow light, and one results in ultraviolet light.

      i. Which lettered change involves the emission of blue light? C

      ii. Which lettered change involves the emission of yellow light? E

      iii. Which lettered change involves the emission of ultraviolet light? A

2. Examine the diagram of an electron cloud below.

Go to http://www.ck12.org/user:Chem12-22/section/The-Bohr-Model-and-the-Quantum-Mechanical-Model/
Scroll down to the section entitle "Probability Patterns" and read that section. What do the dots in the model represent? Be specific.

The model represents an atom with a single electron. The dots represent location where an electron might be found. Where the dots are most dense is the region where one is likely to find an electron. Areas with a more diffuse pattern represent locations where one would be least likely to find an electron.

3. Examine the electron configurations for both platinum and gold in the table below. Write their expected electron configurations based on the rules that you have learned in class and provide an explanation for these two exceptional configurations.

Platinum	Pt*	78	[Xe] $4f^{14}5d^{10}6s^{0}$
Gold	Au*	79	[Xe] $4f^{14}5d^{10}6s^{1}$

Pt [Xe] $6s^2 4f^{14} 5d^8$

In the exceptional configuration of Pt, two electrons from the 6s sublevel are elevated to the 5d sublevel. This completely fills the 5d sublevel.

Au [Xe] $6s^2 4f^{14} 5d^9$

In the exceptional configuration of Au, one electron from the 6s sublevel is elevated to the 5d sublevel. This carpletely fills the 5d sublevel.

4. Why is it that a particle such as an electron can exhibit wave-like properties that cannot be observed in larger objects (Hint: Refer to the work of di Broglie)

Di Broglies equation: $\lambda = \frac{h}{mv}$

Electrons have a very small mass (9.1×10^{-28} g). When (m) in the denominator is very small, the wavelength produced is measureable. Larger particles, with larger mass, produce wavelengths that are undetectable because of their small size.

5. In order for the photoelectric effect to take place in a piece of silver metal, the minimum amount of energy required is 4.7eV. 1eV = 1.6 x 10^{-19}J. Calculate the wavelength of light that corresponds to this energy using the factor label method.

$E = h\nu \qquad c = \lambda \nu$

$\frac{E}{h} = \nu \rightarrow \frac{c}{\nu} = \lambda$

$$\underbrace{\frac{3.0 \times 10^{8}\,m}{s}}_{c} \underbrace{\left(\frac{1}{4.7\,eV}\right)\left(\frac{1\,eV}{1.6 \times 10^{-19}\,J}\right)\left(\frac{6.626 \times 10^{-34}\,J \cdot s}{1}\right)}_{\nu} = \boxed{2.6 \times 10^{-7}\,m}$$